Intramolecular Hydrogen Bonding Involving Hydroxyl Groups in Mono- and Dianions of Diprotic Acids in Acetonitrile and Dimethyl Sulfoxide

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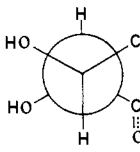
Abstract: Intramolecular hydrogen bonding between OH and COO^- groups has been shown to occur in hydroxyl containing aliphatic straight chain mono- and dicarboxylic acids in AN and DMSO. From estimation of the substituent effect of OH on the α or β carbon atom on the dissociation constant of these acids it was possible to find values of $K^* = [A^{*-}]/[A^-]$, where A^{*-} denotes the anion with intramolecular hydrogen bonding, between OH and COO^-. For monoanions of hydroxy dicarboxylic acids K^* is very much smaller than that of $K' = [\overline{HA^-}]/[HA^-]$, $[\overline{HA^-}]$ denoting the bianion with intramolecular hydrogen bonding between COOH and COO^-. In the monoprotic acids investigated K^* was found to be zero in the amphiprotic solvents water and methanol, while K^* in DMSO is about three times smaller than that in AN. With the monoprotic acids the resolution of acid strength, R, between water and methanol on the one hand and AN and DMSO on the other is: $R_{AN,DMSO} = 1.9R_{W,M}$. Results obtained with hydroxyl containing diprotic acids, with special emphasis on tartaric acids, are discussed and interpreted. As expected, values of K' are smaller in DMSO than in AN. Values of transfer activity coefficients $p{}^{M}\gamma^{W,AN,DMSO}$ of H_2A, HA^-, and A^{2-} have been estimated. The differences in transfer activity coefficients of succinic and tartaric acids between various pairs of solvents, S, are attributed to intermolecular hydrogen bonding between OH and S. The intramolecular hydrogen bond is particularly pronounced in the tartrate ion (A^{2-}) and has a strong stabilizing effect on the ion in AN. Because of intramolecular hydrogen bonding between OH and COO^-, hydration and alcoholation constants of divalent tartrate ion in AN are much smaller than those of succinate.

In the present paper we extend our previous studies of intramolecular hydrogen bonding in monoanions of aliphatic dicarboxylic acids in acetonitrile (AN)[1] and dimethyl sulfoxide (DMSO)[2] to dicarboxylic acids containing hydroxyl groups. These groups are considerably weaker hydrogen bond donors as compared to carboxyl groups but may hydrogen bond intramolecularly to the carboxylate group in monoanions and, to a much greater extent, in the dianions. In a monoanion containing two hydroxyl groups various intramolecularly hydrogen bonded structures are possible, as for example in the *d*-bitartrate ion:



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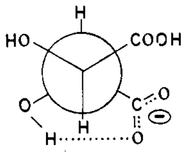

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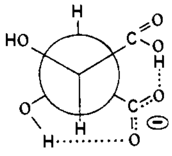

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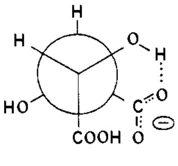

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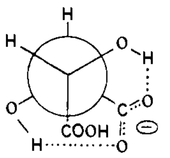


V

Using again[1] a superscript bar to denote intramolecular hydrogen bonding between the COOH and COO^- groups in the monoanion, as was done originally by Westheimer and Benfy,[3] structure I, which does not involve intramolecular hydrogen bonding via the hydroxyl group, is denoted by $\overline{HA^-}$. In addition, either one or two hydroxyl group(s) can hydrogen bond intramolecularly with the COO^- group, while leaving the COOH group free (structures II, IV, and V). Collectively, these species are denoted by an asterisk, i.e., HA^{*-}. There is also a possibility that the COO^- group of the monoanion is intramolecularly hydrogen bonded *both* to an OH and to a COOH group. This species is designated as $\overline{HA^{*-}}$ (structure III). Using the symbols K', K^*, and K'^* for the ratios of the concentrations of the various intramolecularly hydrogen bonded species to the concentration of HA^- we have

$$K' + K^* + K'^* = \{[\overline{HA^-}] + [HA^{*-}] + [\overline{HA^{*-}}]\}/[HA^-] \quad (1)$$

and the first dissociation constant, K_1, of the diprotic acids is:

$$K_1 = a_{H^+}\,\Sigma[HA^-]f_-/[H_2A] = \frac{a_{H^+}\{[HA^-] + [\overline{HA^-}] + [HA^{*-}] + [\overline{HA^{*-}}]\}f_-}{[H_2A]} \quad (2)$$

Denoting the first dissociation constant in the absence of intramolecular hydrogen bonding by K_{H_2A}:

$$K_1 = K_{H_2A}[1 + K' + K^* + K'^*] \quad (3)$$

In similar fashion we obtain for the actual dissociation constant K_E for the monoester of H_2A

$$K_E = K_{HE}(1 + K^*_E) \quad (4)$$

Here, K_{HE} is the dissociation constant expected in the absence of intramolecular hydrogen bonding between OH and COO^-, while K^*_E is

$$K^*_E = [E^{*-}]/[E^-]$$

There is no evidence that the COOR group can act in AN as a hydrogen bond acceptor toward COO^- or OH. By combining eq 3 and 4 and introducing the statistical factor of 2, we obtain:

$$pK_E - pK_1 - \log 2 = \log\left[\frac{1 + K' + K^* + K'^*}{1 + K^*_E}\right] \quad (5)$$

From a comparison of the dissociation constants of glycolic and β-hydroxypropionic acids with those of a group of nonhydroxy substituted monocarboxylic acids in AN and DMSO an estimate was made of the strength of the intramolecular hydrogen bond between the carboxylate and α-hydroxyl or β-hydroxyl groups (K^* values). In similar fashion, values of K^*_E were found from experimentally determined values of pK_E of the monoesters of *d*- and *meso*-tartaric acids. Furthermore, values of pK_1 and pK_2 were estimated in M (methanol, AN, and

Table I. Dissociation Constants of Aliphatic Monocarboxylic Acids in W, M, AN, and DMSO (Conjugation Constants in AN)

Acid	pK_{HA}				$\log K_{HA_2^-}$	$\log K^*$	
	W^a	M	AN	DMSO	AN	AN	DMSO
(1) *n*-Butyric	4.82	$9.69,^c\ 9.9_7^g$	22.7_3^g	12.8_6^g			
(2) Acetic	4.73	$9.6_8,^d\ 9.5_2,^c\ 9.7_0^{e,f}$	22.3^g	12.6^g	3.6^i		
(3) (Et)Me H adipate	4.60	9.8_7^g	22.2_4^h	12.6_4^g	3.6_7^h		
(4) Me H glutarate		9.7_1^g	22.2_6^h	12.4_5^g	3.7_9^h		
(5) Me H succinate	4.49	9.6^g	21.6^h	11.9_1^g	3.7_8^h		
(6) (Et)Me H malonate	3.35	8.8_0^g	20.0^j	10.2_6^g	3.8_7^h		
(7) $ClCH_2COOH$	2.81	$7.8,^f\ 7.8_4^e$	18.8^g	9.0^g			
(8) $NCCH_2COOH$	2.47	$7.5,^f\ 7.5_0^e$	18.0_4^b	8.5_0^b	4.2_8^b		
(9) $BrCH_2BrCHCOOH$	2.17	7.5_3^b	17.1^b	7.1_3^b			
(10) $Cl_2CH{\cdot}COOH$	1.30	$6.3_3,^e\ 6.3,^f\ 6.3_6^b$	15.8^g	6.3_6^b			
(11) $HOCH_2CH_2COOH$	4.51	9.4_2^b	21.0_0^b	11.4_7^b	3.9_3^b	0.5	(0.3)
(12) $HOCH_2COOH$	3.82	$8.8,^f\ 8.7_2^b$	19.3^b	10.2_0^b	3.9_0^b	0.9	(0.5)
(13) *sec*-Bu H *d*-tartrate	3.38^b	8.4_5^b	17.8_3^b	9.0_8^b	4.4_1^b	1.9	1.0
(14) *sec*-Butyl H *meso*-tartrate	3.68^b	8.7^b	17.2_0^b	8.9_7^b		3.0	1.6
(15) Me H tartronate (calcd)^j	(2.44)	(7.8)	(17.0)	(7.8)		(0.9)	(0.5)
(16) Et H oxalate	1.50	6.3_0^h		6.5_2^g			

[a] Compilation of pK_{HA} values in "Determination of Organic Structures by Physical Methods," Vol. 1, E. Braude and F. Nachod, Ed., Academic Press, New York, N.Y., 1955, pp 578–579. [b] This work. [c] J. Juillard, *Bull. Soc. Chim. Fr.*, 1727 (1966). [d] I. Tabagua, *Zh. Fiz. Khim.*, **37**, 1545 (1963). [e] N. Izmailov, V. Chernyi, and L. Spivak, *ibid.*, **37**, 822 (1963). [f] G. Charlot and B. Trémillon, "Chemical Reactions in Solvents and Melts", 1st English ed, Pergamon Press, New York, N.Y., 1969, p 278. [g] Reference 2. [h] Reference 1. [i] Reference 7. [j] Calculated using relation $pK_{HE} = pK_{Me\ H\ malonate} - (pK_{HOAc} - pK_{H\ glycolic})$.

DMSO of the following hydroxyl containing aliphatic dicarboxylic acids: *d*- and *meso*-tartaric, -tartronic, and -tetrahydroxysuccinic acids). Hence, by application of eq 5 values of the sum ($K' + K'^*$) of bitartronate and bitartrates became known, allowing a comparison of the strength of the COOH$\cdots^-$OOC hydrogen bond with that of OH$\cdots^-$OOC in these monoanions.

Transfer activity coefficients, $p^{S_1}\gamma^{S_2}$, between W or M on the one hand, and AN or DMSO on the other, have been estimated for *d*- and *meso*-tartaric acids, their monoanions, and their dianions. Comparison of $p^{M}\gamma^{AN,DMSO}A^{2-}$ of succinate with that of the tartrates allows an evaluation of the strength of intramolecular hydrogen bonding between the OH and COO^- groups in this ion. It is expected that this bond will increase considerably the stability in AN of the tartrate ion as compared to that of the succinate ion. Also, hydration and alcoholation constants of tartaric acid, bitartrate, and tartrate in AN as solvent have been estimated and the effects of OH$\cdots^-$OOC intramolecular hydrogen bonding on hydration and alcoholation of the tartrate ions are discussed.

Results

Dissociation Constants of Mono- and Dicarboxylic Acids in AN, DMSO, and M. Homoconjugation Constants in AN. Dissociation constants of both mono- and dicarboxylic acids in AN, DMSO, and M were estimated from the p*a*H measured with the glass electrode in equimolar mixtures of HA and A^-, H_2A and HA^-, or HA^- and A^{2-}, respectively. In a few systems in AN the ratio of HA to A^- or of H_2A to HA^- was varied to find homoconjugation constants. Plots of p*a*H vs. log $c_{HA}/c_{A^-\text{-f}}$ or log $c_{H_2A}/c_{HA^-\text{-f}}$ are on microfilm (Figure a). All mixtures were prepared by adding 0.91 *M* tetramethylammonium hydroxide in methanol[1] to solutions of the acid. In the estimation of the dissociation constants in AN and DMSO the effect of methanol on p*a*H of the above mixtures was taken into account[1] assuming the methanolation constants of the monovalent anions to be the same as that of monomethylsuccinate anion.[1] As the effect of methanol on p*a*H of the mixtures of HA^- and A^{2-} is considerably larger than that of mixtures of HA and A^-, values of p*a*H were determined at various dilutions of the equimolar mixture ($c = 0.0005$ to -0.003 *M*) and the resulting plots of p*a*H vs. $\sqrt{c}$ extrapolated to $c = 0$. Dissolution of *sec*-butyl *meso*-tartrate ester in AN resulted in formation of a precipitate (or turbidity even at concentrations as low as 2×10^{-3} M) as a result of disproportionation into the diester, which is readily soluble and the sparingly soluble diacid. This was found from alkalimetric titration in water of the precipitate. Therefore, a concentrated solution of the ester in DMSO (0.234 M), which was clear, was introduced into AN with vigorous stirring and then the tetramethylammonium hydroxide solution in M added. The resulting solution was clear. The following p*a*H values were found in AN: $c = 0.001\ 16$, $c_{MeOH} = 0.033$ M, $c_{DMSO} = 0.140$ M, p*a*H = 17.85; 0.000 706, 0.020, 0.085, p*a*H = 17.68; 0.000 360, 0.010, 0.043, p*a*H = 17.47 (p*a*H extrapolated = $pK_{HE} = 17.20$). Activity coefficients of mono- and divalent anions were evaluated in the three solvents using the partially extended Debye-Huckel expression.[2] Values of pK_{HA} in AN, DMSO, and M, and the 1:1 homoconjugation constants in AN of monoprotic acids are listed in Table I. No evidence of dihomoconjugation was found. In Table II are summarized values of pK_1 and pK_2 of diprotic acids in W, AN, and DMSO and the homoconjugation constants of H_2A with ΣHA^- in AN. It is highly improbable that homoconjugation involving ΣHA^- and A^{2-} occurs in AN.

Solubility of Tartaric Acids in Various Solvents. Effect of Additives on Solubility in AN. In Table III the total solubilities (equal to the molecular solubilities) of *d*-, *dl*, and *meso*-tartaric acids in AN are compared with that of succinic acid. The effect of W and M on the solubility of *d*-tartaric acid in AN is represented in the plots of Figure 1. Hydration and methanolation constants in AN derived from these data are in Table IV. As DMSO has been found to form solid solvates with *d*-tartaric acid, the solvation constants of tartaric acid with DMSO in AN cannot be found in this way.

Effect of Additives on paH of Mixtures of Tartaric Acid Species in AN. Solvation Constants. The effect of W, M, or DMSO on p*a*H of *d*- and *meso*-tartaric acid–bitartrate mixtures (curves A) and the corresponding bitartrate-tartrate systems in AN (curves B) are presented in Figure 2. The small

Table II. Dissociation Constants of Diprotic Acids in W, AN, and DMSO (Conjugation Constants in AN)

Acid		pK W	pK AN	pK DMSO	log K_{homo} in AN	log $(K' + K'^*)$ in AN	log $(K' + K'^*)$ in DMSO
Malonic	pK_1	2.8_7	15.3	7.2, 6.6[a]	2.9_5	4.4	2.7_6
	pK_2	5.6_7	30.5	18.5_5, 18.4[a]			
Tartronic	pK_1	2.37	13.8_6	6.8, 6.7[a]	2.2	3.7[b]	1.2[b]
	pK_2	4.74	26.6	14.2_5, 15.2[a]			
Succinic	pK_1	4.20	17.6	9.5, 9.5[a]	2.3_2	3.7	2.1
	pK_2	5.55	29.0	16.7, 16.4[a]			
d-Tartaric[c]	pK_2	3.03	14.9_2	8.1_3, 7.9[a]	1.8	4.5[b]1.6[b]	
	pK_2	4.37	23.3_2	12.2, 11.3[a]			
meso-Tartaric	pK_1	3.22	15.1	7.6_3		4.8[b]	2.6[b]
	pK_2	4.81	24.7	13.7			
Tetrahydroxy-succinic	pK_1		13.7_3	6.8_8	2.7		
	pK_2		26.2	14.0_6			

[a] J. Martin and J. Duperis, *Bull. Soc. Chim. Fr.*, 138 (1968). [b] Calculated from eq 5, assuming K^* is the same in the monoanion as in the anion of the monoester. [c] G. Charlot and B. Tremillon, "Chemical Reactions in Solvents and Melts", Pergamon Press, Elmsford, N.Y., 1969, p 2078, report pK_1 = 7.5 and pK_2 = 9.9 for *d*-tartaric acid in methanol, as compared to 8.2 and 10.4 in the present study. In addition, we find pK_1 = 8.5 and pK_2 = 10.9 for *meso*-tartaric acid in methanol.

Table III. Molar Solubility of Diprotic Acids in Various Solvents

Acid	W	M	AN	DMSO
H_2Suc	0.70	1.28	0.040 4	Solvate
$H_2Tart(d)$	9.28[a]	2.89	0.051 1	
$H_2Tart(d,l)$	1.66, 1.38[a]	0.665	0.004 08	~3.8
$H_2Tart(meso)$	8.35[a,b]		0.040 6	

[a] Mol/kg of water, ref 5. [b] Solid phase monosolvated.

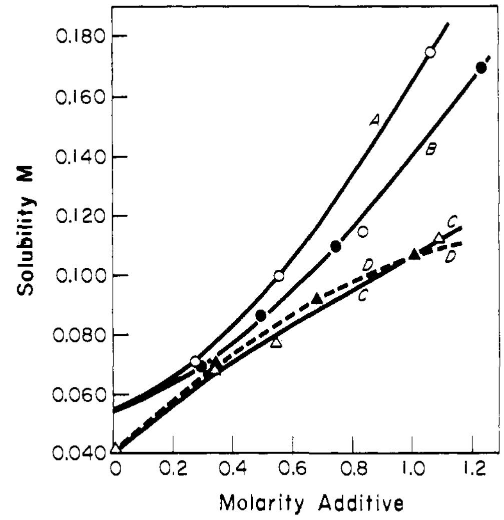


Figure 1. Effect of water of methanol on solubility of succinic and *d*-tartaric acids in acetonitrile. *d*-Tartaric acid with: A, water; B, methanol. Succinic acid with: C, water; D, methanol.

effect of methanol introduced with the tetramethylammonium hydroxide was taken into consideration in the construction of the plots in Figure 2. Alcoholation or hydration constants of $HTart^-$ and $Tart^{2-}$ were evaluated using the data in Figure 2 and the solvation constants of tartaric acid as was done previously for the succinic acid systems.[4] Since DMSO to at least 0.8 M has no effect on paH of 0.0012 M $HTart^-(d)$-$Tart^{2-}$ mixtures in AN, while it increases the paH of H_2Tart-$HTart^-$ mixtures markedly (Figure 2), it is concluded that both bitartrate and tartrate are not appreciably solvated by the solute DMSO. Solvation constants in AN of *d*- and *meso*-tartaric acid species are compared with those of succinic acid species in Table IV.

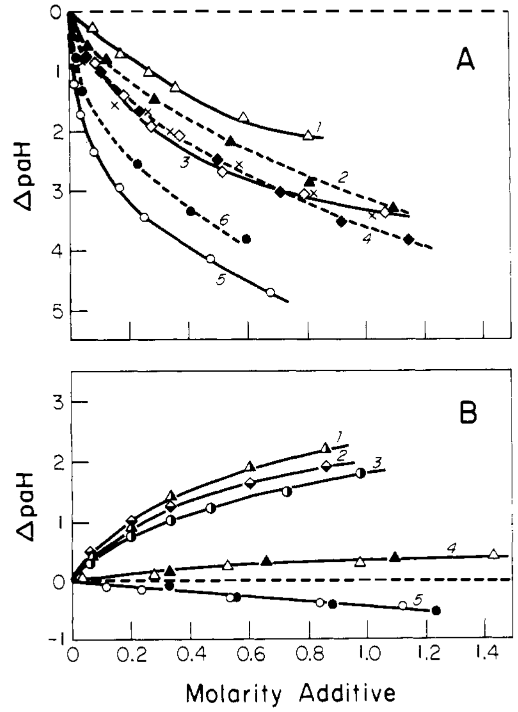


Figure 2. Effect of water, methanol, or DMSO on paH of mixtures of tartaric acid species in AN. A, bitartrate-tartrate system. (1) $c_{HTart^-(d)} = 0.00134$, $c_{Tart^{2-}(d)} = 0.001\,09$, $paH_0 = 23.08$, additive, M; (2) $c_{HTart^-(d)} = 0.000\,536$, $c_{Tart^{2-}(d)} = 0.000\,436$, $paH_0 = 23.13$, additive, W; (3) $c_{HTart^-(meso)} = c_{Tart^{2-}(meso)} = 0.000\,967$, $paH_0 = 24.56$, additive, M; (4) $c_{HTart^-(meso)} = c_{Tart^{2-}(meso)} = 0.000\,387$, $paH_0 = 24.46$, additive, W; (5) $c_{HSuc^-} = 0.001\,29$, $c_{Suc^{2-}} = 0.000\,755$, $paH_0 = 28.3$, additive, M; (6) $c_{HSuc^-} = 0.001\,40$, $c_{Suc^{2-}} = 0.000\,788$, $paH_0 = 28.1$, additive, W; (X) tetrahydroxysuccinate, $c_{HX^-} = c_{X^{2-}} = 0.000\,637$, $paH_0 = 25.8_4$, additive, M. B, tartaric acid-bitartrate system. (1) $c_{H_2Tart(d)} = 0.001\,93$ M, $c_{HTart^-(d)} = 0.001\,14$ M, $paH_0 = 14.68$; (2) $c_{H_2Tart(meso)} = c_{HTart^-(meso)} = 0.000\,967$, $paH_0 = 15.09$, additive, DMSO; (3) $c_{H_2Suc} = c_{HSuc^-} = 0.005\,15$ M, additive, DMSO; (4) $c_{H_2Tart(d)} = 0.001\,93$ M, $c_{HTart^-(d)} = 0.001\,14$ M, additives, W or M; (5) $c_{H_2Suc} = c_{HSuc^-} = 0.004\,06$ M, $paH_0 = 17.49$, additives W or M.

Solubility of Potassium Bitartrates (d and meso) in M and DMSO. The solubility of the tartaric acids is high in M and W (Table III, excepting the *d,l* form), while they form solid solvates in DMSO. In order to find transfer activity coefficients of tartaric acid species involving DMSO, the solubility of potassium *d*- and *meso*-bitartrate was determined in M and DMSO. In water pK^{sp} of potassium *d*-bitartrate is reported at 25 °C as 3.5,[5] and the solubility of the meso salt >0.3 M.

Table IV. Hydration and Alcoholation Constants of Tartrate Species in Acetonitrile

B	$H_2A = H_2Tart(d)$			$H_2A = H_2Tart(meso)$			$H_2A = H_2Suc$		
	W	M	DMSO	W	M	DMSO	W	M	DMSO
$K_{B \cdot H_2A}$	0.67	0.70	0			2.0×10^1	2.0	2.0	1.9×10^1
$K_{2B \cdot H_2A}$	1.8	1.2_5	2.2×10^2			8.6×10^1	0	0	2.8×10^1
$K_{B \cdot HA^-}$	0.6	0.5_2	0	*b*	*b*	0	5.0	5.5	0
$K_{2B \cdot HA^-}$	~0	0	0	*b*	*b*	0	0	0	0
$K_{3B \cdot HA^-}$	0	0	0	*b*	*b*	0	2.8	0	0
$K_{B \cdot A^{2-}}$	(3×10^1)	~0	0	~0	~0	0	5×10^1	1.5×10^3	0
$K_{2B \cdot A^{2-}}$	$(2._5 \times 10^2)$	1.5×10^2	0	0.5×10^3	0.8×10^3	0	1.7×10^4	2.4×10^4	0
$K_{4B \cdot A^{2-}}$	1.6×10^3	*a*	0	3.2×10^3	1.2×10^3	0	3.2×10^5	2.2×10^5	0

[a] Higher methanolate than dimethanolate present, but constant is small and highly uncertain. [b] Hydration or methanolation of monoanion assumed same as that of $HTart^-(d)$.

Table V. Solubility of Potassium Hydrogen Tartrates in Various Solvents

K bitartrate salt taken	Solvent	log a_{K^+}[a]	f_{K^+}	$[K^+]$[b]	Conductivity of satd. soln., $ohm^{-1} cm^{-1}$	Λ_{KHA}	*d*	pK^{sp}[c]
$HTart^-(d)$	DMSO	$-3.0_4, -3.1_4$	0.92	$1.0_0 \times 10^{-3}$	2.08×10^{-5}	24.7		6.3
$HTart^-(meso)$	DMSO	$-2.1_6, -2.1_7$	0.78	8.8×10^{-3}	1.56×10^{-4}	$17._7$		4.3
$HTart^-(d)$	MeOH	$-4.1_5, -4.1_8$	0.96	0.69×10^{-4}	7.1×10^{-6}	102		8.3_6
$HTart^-(meso)$	MeOH	$-3.1_8, 3.2_2$	0.89	6.9×10^{-4}	$5.3_5 \times 10^{-5}$	80	8.3×20^{-4}	6.4

[a] In saturated solution, cation electrode uncorrected; corrected for electrode response to H^+. [b] Calculated from a_{K^+} found with specific ion electrode and f_{K^+} calculated from the partially extended Debye–Huckel relation. [c] Calculated assuming $a_{K^+} = a_{HTart^-}$. [d] Total solubility determined by evaporating solvent and titrating residue alkalimetrically in water.

Hence, the solubility method is not suitable for estimation of transfer activity coefficients of *meso*-tartrate species from water to S. Solubility data of potassium bitartrate (d and meso forms) in M and DMSO are summarized in Table V. The third column lists the values of $-pa_{K^+}$ in the saturated solutions obtained with the cation specific electrode. In these solutions the effect of H^+ on electrode potential is small and has been estimated in methanol in 6.1×10^{-5} to 6.1×10^{-4} M potassium perchlorate at pH 7.29–9.7 in 3,4-dimethyl- or 3,5-dinitrobenzoic acid buffers and in DMSO in 1.6×10^{-3} to 1.2×10^{-2} M potassium perchlorate buffered at pH 7.4 to 10.7 in *m*-bromobenzoic acid systems. Data are available upon request. Using the conductivity data of the saturated potassium bitartrate solutions in the sixth column in Table V and the value of $[K^+]$ (fifth column, Table V), derived from a_{K^+} found with the cation electrode, values of Λ_{KHA} are found and are listed in the seventh column.

From $\lambda_{0K^+} = 52.4$ in M[6] and 13.9 in DMSO,[6] we obtain $\lambda_{0HTart^-(d)} = (52)$ in M and $(12._3)$ in DMSO and $\lambda_{0HTart^-(meso)} = (34)$ in M and $(8._1)$ in DMSO. These values are low, but are reasonable, considering errors introduced in estimation of $[K^+]_{sat}$. The total solubility in methanol for both bitartrate salts (column 8) indicates that they are extensively dissociated in their saturated solutions.

A check for disproportionation of the bitartrate salts into the sparingly soluble normal tartrates and the very soluble diacid was made by estimating pH visually with indicators. Using dibromothymolbenzein and bromcresol green, agreement within 0.5 pH unit with the calculated value, $(pK_1 + pK_2)/2$, was obtained. Values of p*K* in DMSO[7] and in M[8] were taken from previous publications. The glass electrode would be subject to considerable potassium ion error in these solutions.

Discussion

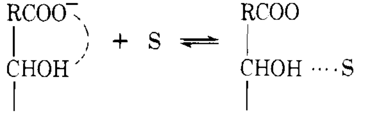
Dissociation Constants of Monoprotic Acids. Estimated Values of K^*. Using the data in Table I, a plot of pK^d_{HA} values of the various monoprotic acids in AN vs. those in M was constructed (Figure 3). The nonhydroxy acids (solid circles) yield the following least-squares linear relation:

$$(pK_{HA})_{AN} = 1.86(pK_{HA})_M + 4.06$$

Acids containing α and/or β hydroxy substituents (open circles) are stronger in AN than expected from the straight line. This deviation, the magnitude of which was taken equal to log $(K^* + 1)$ (cf. eq 4), is attributed to intramolecular hydrogen bonding between the OH and COO^- groups in the anion. Values of log K^* are listed in Table I. A plot of $(pK_{HA})_{DMSO}$ vs. $(pK_{HA})_M$, also in Figure 3, has a similar appearance as that of $(pK_{HA})_{AN}$ vs. $(pK_{HA})_M$ and yields the relation

$$(pK_{HA})_{DMSO} = 1.76(pK_{HA})_M - 4.72$$

for the nonhydroxy acids. As expected, the values of K^* in DMSO are smaller than in AN (Table I). The equilibrium



is displaced to the right to a greater extent in the much stronger hydrogen bond acceptor solvent DMSO than in AN. In amphiprotic solvents (W, M) COO^- and OH are hydrogen bonded to the solvents. The values of log K^* of the *d*-tartrate monoester in AN and in DMSO are found to be close to the sum of log K^* values of glycolate and β-hydroxypropionate (additive free energies of formation of the $-OH \cdots O^-OC-$ hydrogen bonds), while the values of log K^* of the *meso*-tartrate monoester in these solvents are considerably larger. The acid strengthening effect of the two OH groups of the *d*-tartrate ester has been estimated to be 3.8 and 2.8 units in AN and DMSO, respectively, from the difference in pK_E between monomethyl succinate and tartrate monoester. Part of this difference is due to the electron withdrawing inductive effect of the OH groups. Approximately one-half of the difference, $pK_{MeHSuc} - pK_{RHTart(d)}$, in AN and one-third of this difference in DMSO is attributed to intramolecular hydrogen

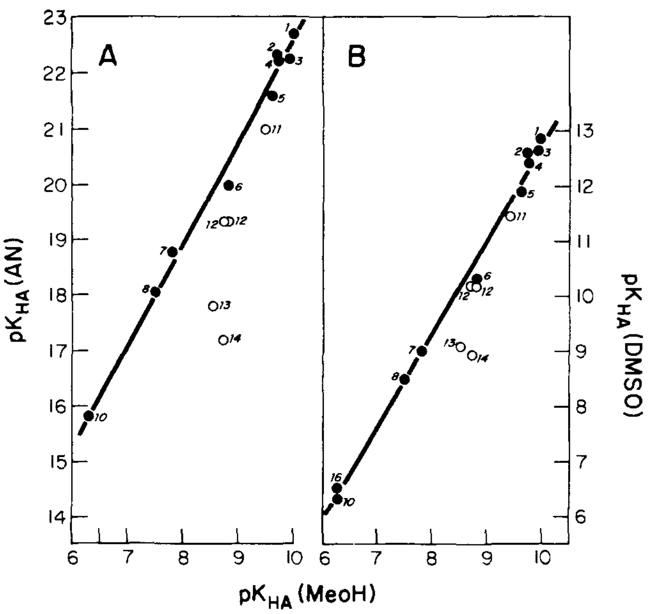


Figure 3. Plots of $\mathrm{p}K(\mathrm{HA})_{\mathrm{AN}}$ (line A) or $\mathrm{p}K(\mathrm{HA})_{\mathrm{DMSO}}$ (line B) vs. $\mathrm{p}K(\mathrm{HA})_{\mathrm{M}}$ for aliphatic monocarboxylic acids. Left-hand ordinate for A, right-hand ordinate for B. Closed circles, nonhydroxy; open circles, hydroxy acids. Numbers are same as in Table I.

bonding between the OH and $\mathrm{COO^-}$ groups (i.e., $\log K^*$ values of the *d*-tartrate monoester in Table I).

It is of interest to note that a plot in Figure 4, line B, of $(\mathrm{p}K_{\mathrm{HA}})_{\mathrm{M}}$ vs. $(\mathrm{p}K_{\mathrm{HA}})_{\mathrm{W}}$ of the monoprotic acids in Table I yields the following straight line of virtually unit slope

$$(\mathrm{p}K_{\mathrm{HA}})_{\mathrm{M}} = 1.03(\mathrm{p}K_{\mathrm{HA}})_{\mathrm{W}} + 5.00$$

This means that (a) there is no resolution of acid strength of substituted acetic acids between W and M, (b) that the inductive effect of an α-OH group (and also a β-OH group) is the same in both solvents, and (c) that there is no intramolecular hydrogen bonding between the OH and $\mathrm{COO^-}$ groups in W and M, both solvents being strong hydrogen bond donors toward the carboxylate and weak hydrogen bond acceptors toward OH groups.

Dissociation Constants of Diprotic Acids. It was found previously[2] that the monoanions of the homologous series of oxalic acid having three or more $\mathrm{CH_2}$ groups in the chain are not intramolecularly hydrogen bonded to any appreciable extent in W and M. Values of $\mathrm{p}K_1$ of these acids lie on the linear plot of $(\mathrm{p}K_{\mathrm{HA}})_{\mathrm{M}}$ vs. $(\mathrm{p}K_{\mathrm{HA}})_{\mathrm{W}}$ of the aliphatic monocarboxylic acids (Figure 4, line B). The differences in values of $\mathrm{p}K_1$ and $\mathrm{p}K_{\mathrm{HE}}$ in W and M (footnote *c* in Table II) of *d*- or *meso*-tartaric acids and their *sec*-butyl monoesters in the present study are 0.35 and 0.4 in W and 0.4 and 0.2 in M, respectively (Tables I and II), indicating little or no intramolecular hydrogen bonding in the bitartrate monoanions in these protic solvents. The monoanions of the homologous series of oxalic acid having up to seven $\mathrm{CH_2}$ groups in the chain are intramolecularly hydrogen bonded in AN and DMSO.[2]

The values of $\log K^*$ of anions of hydroxymonocarboxylic acids are considerably smaller than those of $\log K'$ of monoanions of nonhydroxydicarboxylic acids of the same chain length, the difference being about 3 units in AN and 2 in DMSO (Tables I and II). This is expected, as the carboxyl group is a much stronger hydrogen bond donor than the hydroxyl group. From ir spectra to be presented in a subsequent paper valuable information was obtained on the structure of monoanions of dicarboxylic acids. The *d*-bitartrate and bitartronate ions in AN appear to be present primarily as $\overline{\mathrm{HA}}^{*-}$ (structure III). This form possesses a strong $-\mathrm{COOH}\cdots\mathrm{{}^-OOC}-$ intramolecular hydrogen bond. A solution of these monoanions in a weakly basic aprotic solvent, like AN, is essentially an equilibrium mixture of structures I and III, structures II, IV, and V apparently being nonexistent. Values of $\log (K' + K'^*)$ listed in Table II were calculated using eq 5, assuming $K^* = K^*_{\mathrm{E}}$ and substituting values of K^*_{E} of the corresponding monoester from Table I.

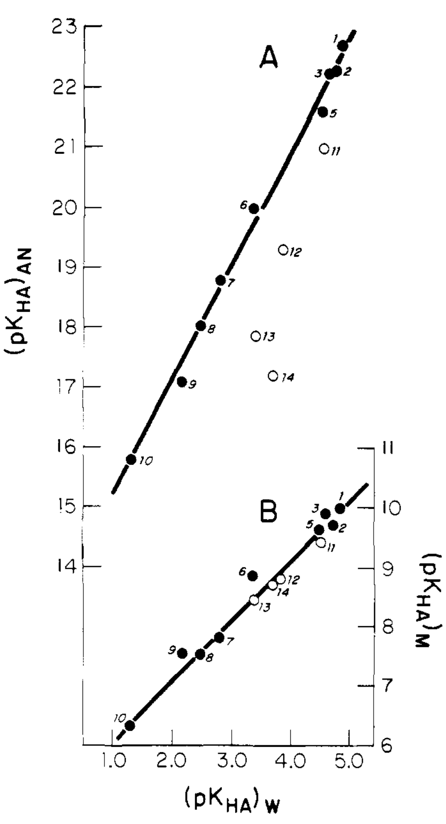


Figure 4. Plots of $\mathrm{p}K(\mathrm{HA})_{\mathrm{AN}}$ (line A) or $\mathrm{p}K(\mathrm{HA})_{\mathrm{M}}$ (line B) vs. $\mathrm{p}K(\mathrm{HA})_{\mathrm{W}}$ for monocarboxylic acids. Left-hand ordinate for A, right-hand ordinate for B. Symbols are the same as in Figure 3 and the numbers the same as in Table I.

Conversion of structure I to III involves formation of a single $\alpha\text{-OH}\cdots\mathrm{{}^-OOC}$ intramolecular hydrogen bond, the strength of which can be taken approximately equal to that in glycolate.[9] Hence $[\overline{\mathrm{HA}}^{*-}]/[\overline{\mathrm{HA}}^{-}] = K^*$ glycolate. Combining the values of $\log K^*$ glycolate in Table I with those of $\log (K' + K'^*)$ in Table II permits the estimation of K' and K'^* individually. The following values of $\log K'$ are found: 3.6, 3.9, and 2.8 for *d*-bitartrate, *meso*-bitartrate, and bitartronate in AN, respectively. Corresponding $\log K'$ values in DMSO are 1.0, 2.0, and 0.6. The values of $\log K'$ of bitartrates generally are close to those of bisuccinate, but those of bitartronate are considerably smaller than those of bimalonate. It appears that the OH substituents in bitartrate have little effect on the strength of the $-\mathrm{COOH}\cdots\mathrm{{}^-OOC}-$ intramolecular hydrogen bond in $\overline{\mathrm{HA}}^-$ (structure I). For reasons mentioned in a previous paper,[2] values of $\log K'$ of the homologous series of oxalic acids are 1.7 units smaller in DMSO than in AN. On the basis of the assumptions made in the evaluation of K', a similar difference between $\log K'$ values in AN and DMSO is encountered for the monoanions of the hydroxy dicarboxylic acids.

From the data in Tables I and II it is evident that the intramolecular hydrogen bonds $-\mathrm{OH}\cdots\mathrm{{}^-OOC}-$ in structures II–V, particularly in structure III, have hardly any effect on the homoconjugation constants of the various mono- and diprotic acids. Furthermore, homoconjugate species involving *inter*molecular $-\mathrm{OH}\cdots\mathrm{{}^-OOC}-$ hydrogen bonding in the diprotic monoanions can be considered as negligible.

Table VI. Transfer Activity Coefficients of Tartaric Acid Species

	$p^{AN}\gamma^{W}$	$p^{AN}\gamma^{DMSO}$	$p^{AN}\gamma^{M}$	$p^{W}\gamma^{DMSO}$	$p^{M}\gamma^{DMSO}$
H_2Suc	−1.2	(−1.9)	−1.5	−0.7	(−0.4)
$H_2Tart(d,l)$	-2.6^a		-2.2^a		
$H_2Tart(d)$		-6.1^d	$(-1.7)^a$	$(-3.9)^a$	-3.3^c
			-2.8^e	-3.5^c	
$H_2Tart(meso)$					-2.5^c
$HSuc^-$	−6.5	+1.4	−3.8	+7.9	+5.2
$HTart^-(d)$	-6.4^g	-1.5^f	-3.3^f	$+4.9^b$	$+1.8^b$
$HTart^-$-(*meso*)					$+1.8^b$
Suc^{2-}	−21.9	+0.5	−15.1	+22.4	+15.6
$Tart^{2-}(d)$	−17.2	−1.2	−10.0	+16.0	+9.0
$Tart^2(meso)$					+9.8

[a] From solubility of diacid. [b] From solubility of KHTart and $p\gamma K^+$. [c] From $p\gamma HTart^-$ calculated from *b*, $p\gamma H^+$, and pK_1. [d] $p^{AN}\gamma^{DMSO}H_2Tart(d) = p^W\gamma^{DMSO}H_2Tart(d) + p^{AN}\gamma^W H_2Tart(d,l)$, using the value of $p^W\gamma^{DMSO}H_2Tart(d)$ not in parentheses and assuming $\gamma H_2Tart(d,l) = \gamma H_2Tart(d)$. [e] $p^{AN}\gamma^M H_2Tart(d) = p^{AN}\gamma^{DMSO}H_2Tart(d) - p^M\gamma^{DMSO}H_2Tart(d)$, i.e., *c* and *d*. [f] From values of $p^{AN}\gamma^{M,DMSO}H_2Tart$, pK_1, and $p^{AN}\gamma^{M,DMSO}H^+$ using $p^{AN}\gamma^{M,DMSO}H_2Tart$ values not in parentheses. [g] $p^{AN}\gamma^W HTart^-(d) = p^{AN}\gamma^{DMSO}HTart^-(d) - p^W\gamma^{DMSO}HTart\ (d)$.

Transfer Activity Coefficients, γ. Values of $p^{AN}\gamma^S$ of the various species are listed in Table VI. Transfer activity coefficients of ions are based on the tetraphenylborate assumption. One-half of the difference in transfer activity coefficient of *d*-tartaric and succinic acids between AN and S, $(p^{AN}\gamma^S H_2Tart - p^{AN}\gamma^S H_2Suc)/2$, is regarded as the difference in free energy of solvation (in log *K* units) per OH group in tartaric acid in AN and S. This has been found equal to −2.1, −0.6, and −0.7 (Table VI), S being DMSO, M, and W, respectively. These values are comparable to $p^{AN}\gamma^S(H_a)$, which is the transfer energy, γ, of the intermolecular hydrogen bond between a carboxyl group in AN and the solvent, S, as acceptor. Values of $p^{AN}\gamma^S(H_a)$ have previously been reported[10] as −1.9, −1.4, and −0.8, S being DMSO, M, and W. Regarding the nonhydrogen bonded contribution of the OH group to $p^{AN}\gamma^S(H_a)$ as small and constant for the various solvents, it is concluded that the order of strength of the intermolecular hydrogen bond between COOH and S parallels that between OH and S. In other words, we find the *order* of strength of the intermolecular hydrogen bond between a solvent acceptor and a donor independent of the strength of the hydrogen bond donor.

Hydration and Alcoholation of Tartrate Species in AN. In accordance with the transfer activity coefficients in Table VI the overall solvation with DMSO, M, and W of the tartaric acids in AN (Figure 1, Table IV) is considerably greater than that of succinic acid. From the constants in Table IV it is apparent that the hydration and methanolation of the dianions are in the order: succinate > *meso*-tartrate > *d*-tartrate, while DMSO as a solute does not solvate any of these ions. In agreement with this order is the observation of Mathieson and Conway[11] that the apparent molal volumes and molal adiabatic compressibilities in aqueous solution are larger for the *meso*- than for the *d*-tartrate. Apparently, water and methanol in dilute solutions in AN are weaker hydrogen bond donors than the OH group in the intramolecular bond with COO^-.

Experimental Section

Chemicals. Solvents. Acetonitrile, methanol, and dimethyl sulfoxide were purified and stored as described previously.[1] The preparation of 0.9 *M* tetramethylammonium hydroxide in methanol has been described[1] elsewhere.

Acids. Monochloro-, α,β-dibromo-, and dichloroacetic acids were Eastman Kodak White Label products. The dichloro acid was distilled at 20 mm of pressure, while the others were used as received. Cyanoacetic, glycolic, and β-hydroxypropionic acids were from Aldrich Co. The latter two were not further purified, while the former was recrystallized from chloroform. Assay of the above acids by alkalimetric titration in water with the exception of β-hydroxypropionic acid (which contained 6.8 mol % bimolecular esterification product) was 99.5 ± 0.4%. Mono *sec*-butyl esters of *d*- and *meso*-tartaric acids were prepared by refluxing an equimolar mixture of the alcohol with the diacid in the presence of a few drops of concentrated sulfuric acid as catalyst.[12] The monoester was used without further purification as recommended.[12] Recrystallization from anhydrous ether resulted in disproportionation of these esters into the insoluble diacid and soluble diester. Assay by direct alkalimetric titration in water and from back-titration with standard HCl to a bromthymol blue end point after saponification by boiling ½ h with a known amount of sodium hydroxide yielded the following results: RHA, 97.5 wt %; H_2A, 2.5%; no detectable diester in the *d*-tartrate monoester; RHA, 97.7 wt %; R_2A, 2.3%; no detectable diacid in the *meso*-tartrate monoester. Aldrich Co. tartronic and *d*-tartaric acids were used as received, while tetrahydroxysuccinic, *meso*-tartaric acid monohydrate, and *d,l*-tartaric acid hydrate were recrystallized from AN or AN–ethanol mixtures and dried at 60 °C at atmospheric pressure. Assay of the latter two acids corresponded to the anhydrous form.

Salts. Tetramethylammonium glycolate, bitartrate, and *d*-tartrate were prepared in the same way as tetraethylammonium salts of diprotic acids.[1]

Crystal Solvates of Acids and Salts. Using a procedure described in a previous paper[10] the following were found not to form solvates: *d,l*-tartaric acid with water or DMSO and potassium *d*-bitartrate with DMSO. *d*-Tartaric acid formed the solvate $H_2Tart{\cdot}0.64DMSO$.

Techniques. The procedure used in determining the solubilities of tartaric acids in the various solvents has been described.[2] The total solubility of potassium bitartrates was estimated in methanol in a similar way. The potentiometric cell, reference electrode, and pH meter were the same as used previously.[13] Calibration of the Beckman No. 39137 cationic glass electrode in methanol was carried out in saturated solutions of potassium bromide, chloride, and perchlorate, the values of pK_{sp} having previously been reported.[14] In DMSO dilute potassium perchlorate solutions were used for calibration. In both solvents the electrode response was 57 ± 1 mV/p*aK* unit in the p*aK* range covered in the present work.

Acknowledgment. We thank the National Science Foundation for Grants GP-20605 and MPS-7506591 in support of this work.

Supplementary Material Available: Plots of p*a*H in AN vs. log c_{HA}/c_{A^-} for substituted acetic acids or vs. log c_{H_2A}/c_{HA^-} for aliphatic hydroxy dicarboxylic acids (2 pages). Ordering information is given on any current masthead page.

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